

# **TOPIC 17 – EQUILIBRIUM**

## **17.1: LIQUID–VAPOR**

### **EQUILIBRIUM**

IB Chemistry

T17D04



# 17.1 **Liquid/vapor Equilibrium - 2 hours**

- 17.1.1 Describe the equilibrium established between a liquid and its own vapor and how it is affected by temperature changes. (2)
- 17.1.2 Sketch graphs showing the relationship between vapor pressure and temperature and explain them in terms of the kinetic theory. (3)
- 17.1.3 State and explain the relationship between enthalpy of vaporization, boiling point and intermolecular forces. (3)



# Intro - Refrigeration

- The invention of refrigeration has allowed for the storage and transport of perishable materials
- Food lasts longer
- Food can be shipped or moved to remote places where it may not be available



- SS Dunedin (1882)
- The New Zealand sailing vessel was fitted with a refrigeration unit.
- Led to a meat and dairy boom in Australia, New Zealand, and South America



# Intro - Refrigeration

- The key behind refrigeration is that evaporation ( $\Delta H_{\text{vap}}$ ) is an endothermic process
- In a refrigerator, a liquid with a low boiling point circulates around a circuit of pipes
  - ENDO - Liquid evaporates in the pipes, taking in heat energy from the air inside the refrigerator
  - EXO - refrigerator compresses the vaporized gas (which is hot), and as it flows through the pipes at the rear of the refrigerator it condenses back to a liquid, heating up the air behind the refrigerator
- Refrigerator transfers energy from the inside of the fridge to the air in the room



# Intro - Refrigeration

- Evaporation is
  - Particles at the liquid surface
  - $KE > \text{intermolecular forces}$
  - Can occur at any temperature
  - If in a closed system can enter physical equilibria

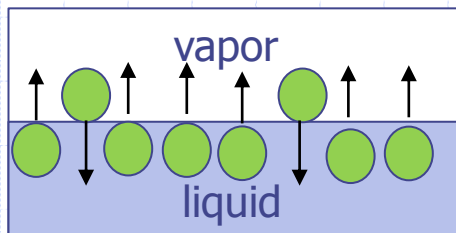


# 17.1.1 – Liquid-Vapor Equilibrium

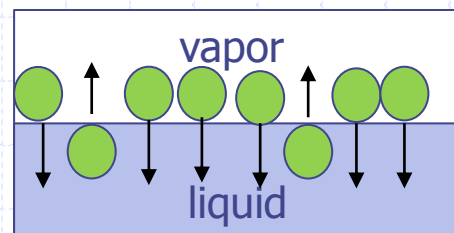
- 17.1.1 Describe the equilibrium established between a liquid and its own vapor and how it is affected by temperature changes. (2)
- The equilibrium between vapor and liquid:



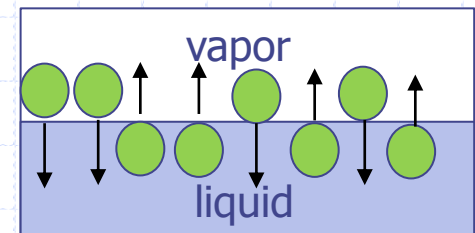
- The position of equilibrium depends on the liquid being used, and the temperature



Evaporation



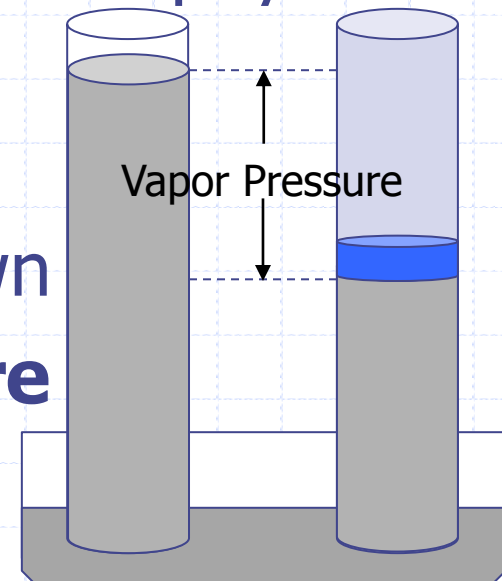
Condensation



Equilibrium

# 17.1.1 – Measuring Vapor Pressure

- Can be explored in an evacuated chamber filled with mercury like a barometer
  - Liquid is injected into evacuated space
  - Liquid floats on top of Hg
  - Liquid evaporates and reaches dynamic physical equilibria
  - The value attained will be constant for a given temperature and known as the **saturated vapor pressure**



$$\text{Rate}_{\text{evaporation}} = \text{Rate}_{\text{condensation}}$$



## 17.1.1 – Relative Vapor Pressure

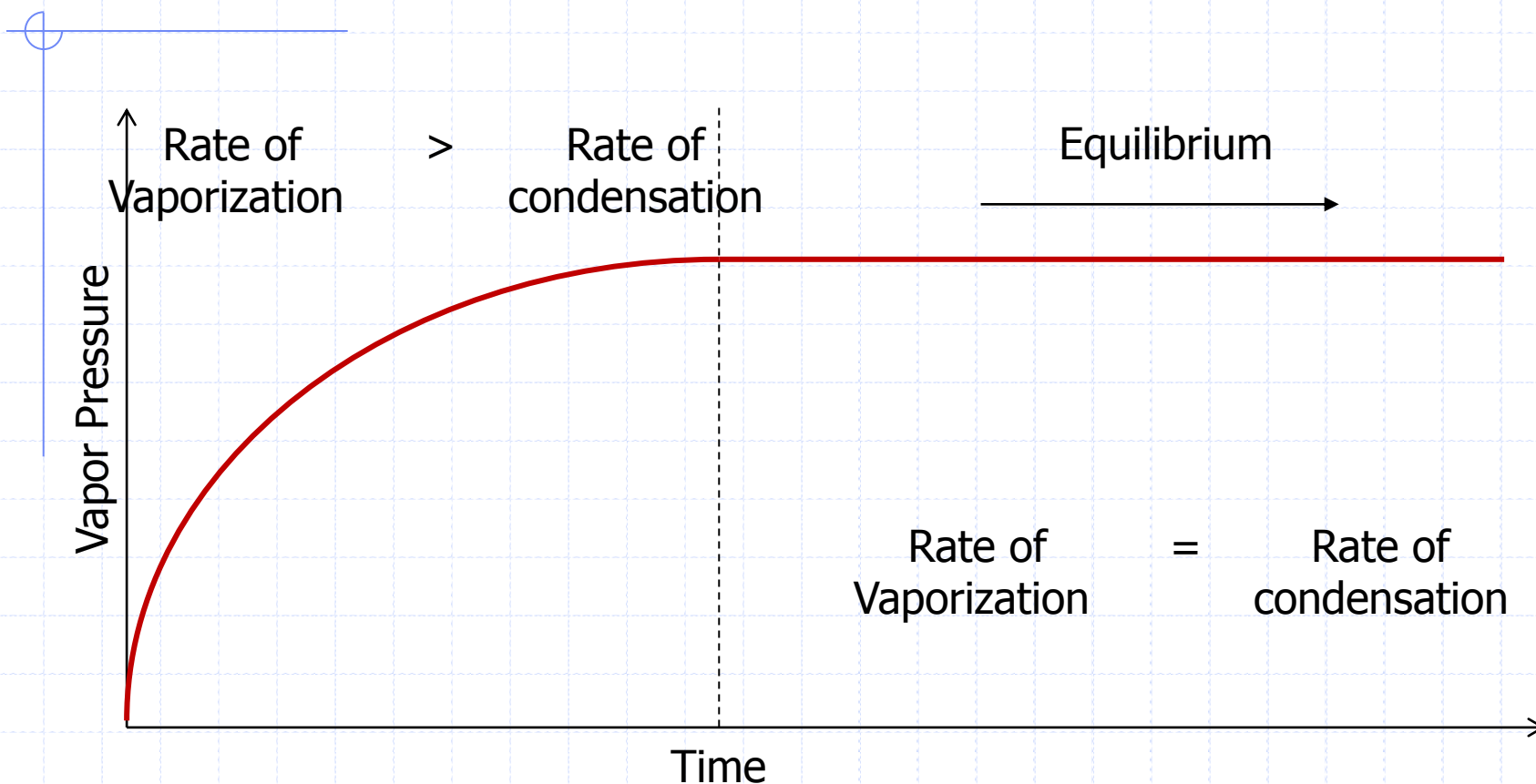
- The table below gives some values of saturated vapor pressure for certain liquids at 298K
- The unit of Pa (SI) used to be presented as mmHg
- What factors cause the differences?
- Why is it crucial that the value for Hg be so much lower than the others?

The **saturated vapor pressure** of a liquid is the pressure exerted by its vapor when two phases are in dynamic equilibrium in a closed system at a given temperature

Liquid	Saturated Vapor Pressure (Pa)
Water	3170
Methanol	16,350
Ethanol	7850
Benzene	12,600
Mercury	0.226



# 17.1.1 – Vapor Pressure / Time



This illustration depicts what happens when a liquid is introduced to a closed container

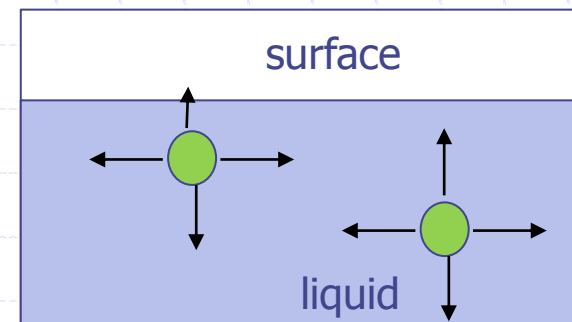
## 17.1.1 – Saturated Vapor Pressure

- The **saturated vapor pressure** of a liquid is the pressure exerted by its vapor when two phases are in dynamic equilibrium in a closed system at a given temperature. (given at 298K)
  - If evaporation is very slow, the pressure of the vapor does not need to be very high for the condensation rate ( $\propto$  pressure) to match it
    - Low SVP is a sign that molecules are leaving the surface at a relatively low rate
  - If evaporation is rapid, the vapor pressure will reach a high value before condensation can match
    - High SVP is a sign that molecules are leaving the liquid surface at a greater rate



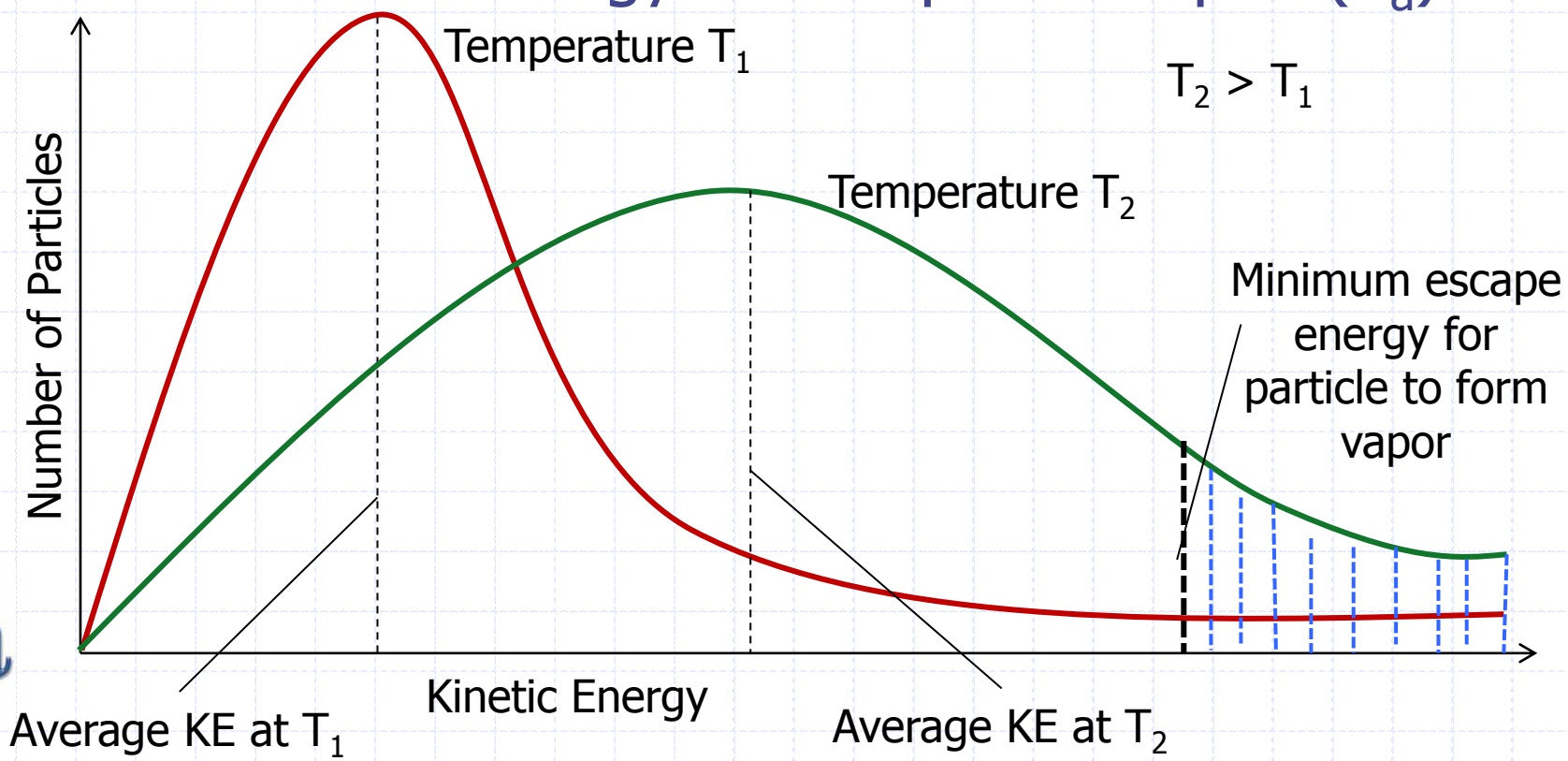
## 17.1.2 – Graphs for Vapor Pressure

- 17.1.2 Sketch graphs showing the relationship between vapor pressure and temperature and explain them in terms of the kinetic theory. (3)
- Evaporation takes place on the surface of the liquid because the molecules there are less strongly bonded than those in the body of the liquid



## 17.1.2 – Max/Boltz Distribution

- With increasing temperature, the molecules in a warm liquid (on average) have more KE
  - Move at a higher average speed
  - Have sufficient energy to escape the liquid ( $E_a$ )

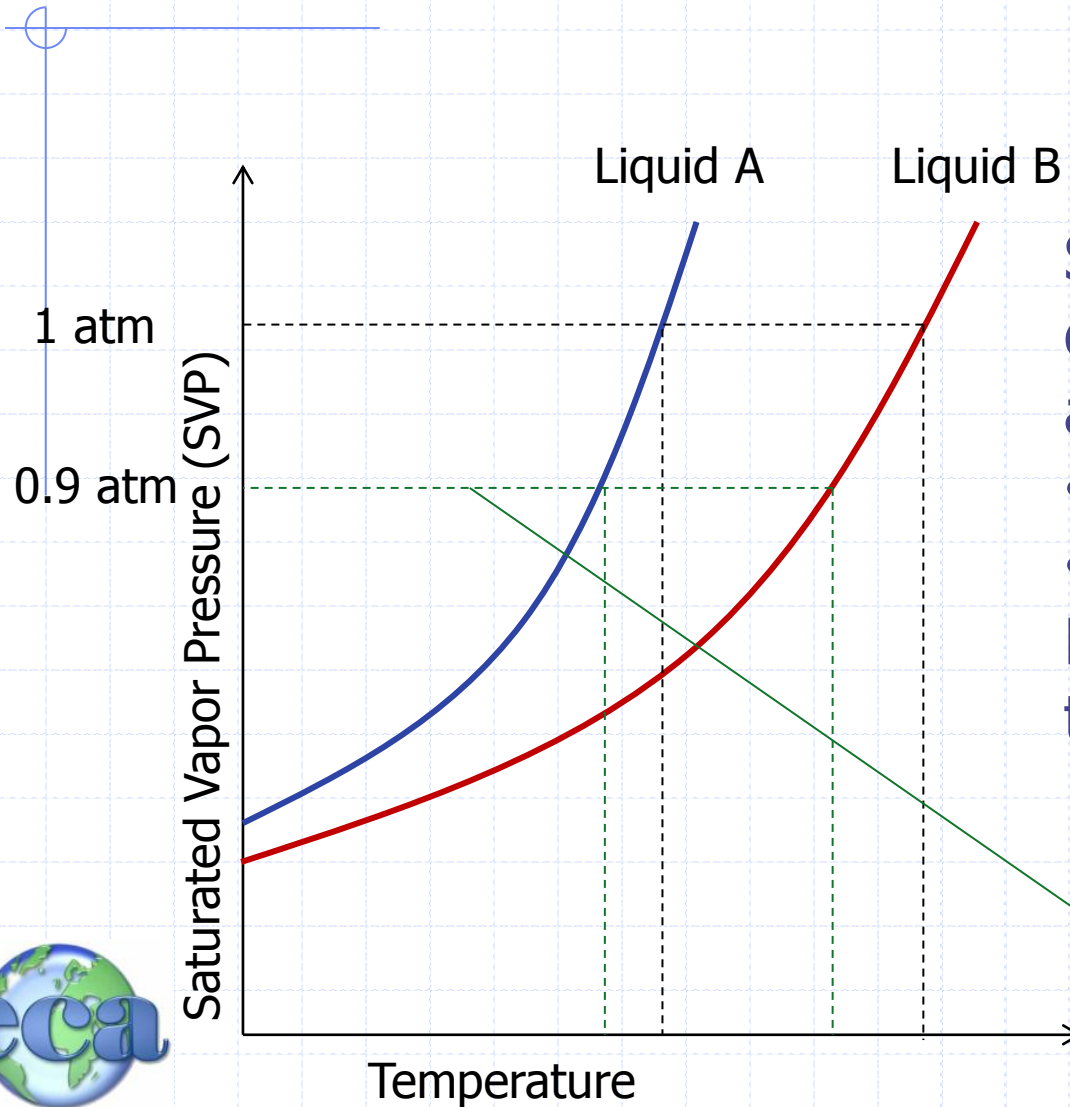


## 17.1.2 – Boiling Point

- As a liquid is heated, its vapor pressure increases.
- When the temperature of the liquid is raised to the point of which vapor pressure = external pressure, then vaporization can occur.
  - At this point, bubbles (boiling) begins to appear and vaporization is much quicker as it's not just at the surface but throughout the liquid
  - The temperature at which this occurs is called the '**boiling point.**'
  - This explains why boiling point is different when containers are not sealed or liquid is at a different elevation. It's because of the atmospheric pressure



# 17.1.2 – Vapor Pressure vs Temp



Saturated Vapor Pressure curves for two liquids A and B, where

- A is ethanol (BP 78°C)
- B is water (BP 100°C)

Ethanol is more volatile than water

Explains why water boils at a lower temperature in places like the summit of Mt. Everest (at 65°C)



## 17.1.3 – $\Delta H_{\text{vap}}$ , B.P., and Intermolecular Forces

- 17.1.3 State and explain the relationship between enthalpy of vaporization, boiling point and intermolecular forces. (3)
- Evaporation is an endothermic process (put a volatile liquid in the palm of your hand and feel the cooling effect of evaporation)
  - Heat energy from the surroundings (your hand) is used to overcome the intermolecular forces keeping the liquid intact
  - THIS DOES NOT BREAK BONDS!! JUST A DISTURBANCE IN INTERMOLECULAR FORCES, REMEMBER THIS IS A PHYSICAL CHANGE!!!!





## 17.1.3 – Enthalpy of Vaporization

- The energy involved in evaporation is referred to as **enthalpy of vaporization** ( $\Delta H_{\text{vap}}$ )
  - The amount of energy required to convert one mole of pure liquid to one mole of the gas at its normal boiling point
- A volatile liquid = high vapor pressure = weak intermolecular forces = low  $\Delta H_{\text{vap}}$
- In contrast, liquids with strong hydrogen bonding or large molecules with lots of van der Waals forces have much greater intermolecular forces and hence low vapor pressure and high  $\Delta H_{\text{vap}}$

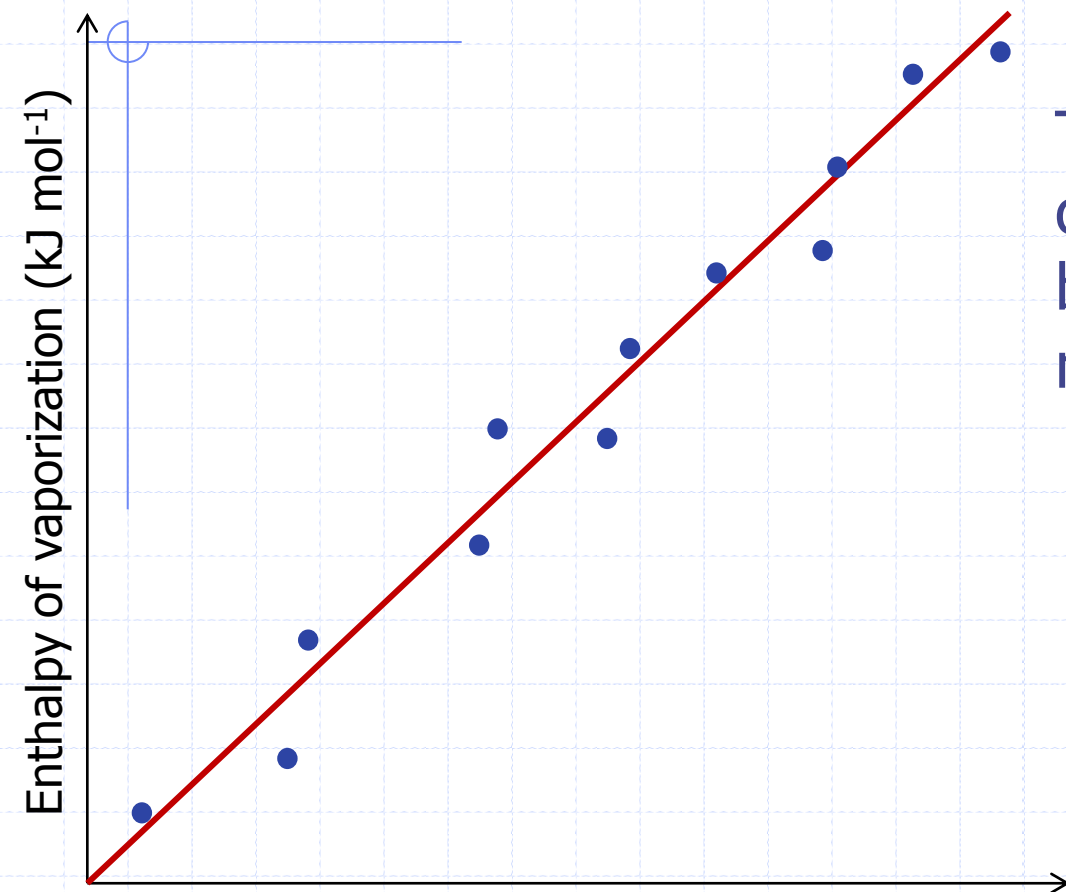


## 17.1.3 – Intermolecular Forces

- Remember three common intermolecular forces from chapter 04
  - van der Waals' forces (temporary induced dipoles)
  - Dipole-dipole interactions (b/w permanent dipoles)
  - Hydrogen bonds (H on N, O, or F to N, O, or F)
- What is the strongest for each compound below:
  - Ethanal -  $\text{CH}_3\text{CHO}$
  - Ethanol -  $\text{CH}_3\text{CH}_2\text{OH}$
  - 2-methylpropane -  $(\text{CH}_3)_2\text{CHCH}_3$
  - Methoxymethane -  $\text{CH}_3\text{OCH}_3$
- Put the first three in order of increasing BP:



## 17.1.3 – BP vs $\Delta H_{\text{vap}}$



There is a very strong correlation between the boiling points and  $\Delta H_{\text{vap}}$  for many different molecules



Boiling Point (K)

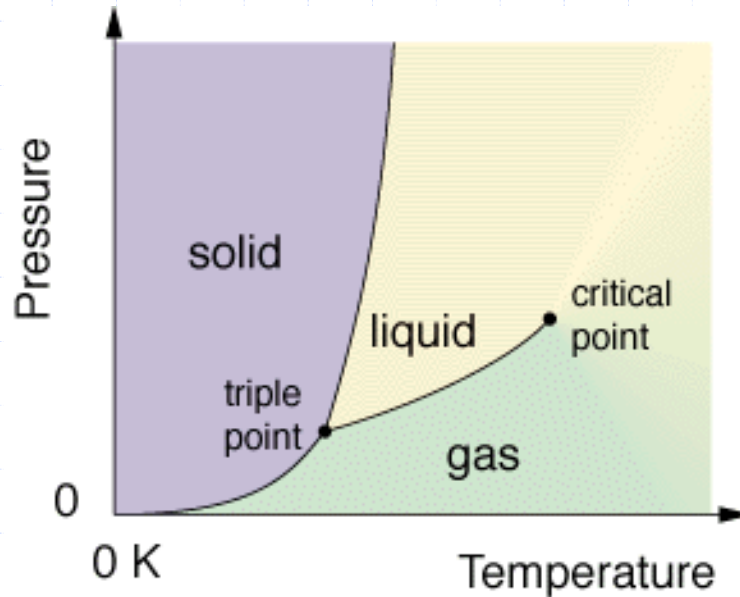
## 17.1.3 - Phases

- A **phase** is defined as a homogeneous part of a system which is chemically and physically uniform throughout
- Phases are separated by phase boundaries
- A number of foods consist of one phase dispersed through another, these are known as **colloids** or **dispersed phases**
- A gas in contact with a liquid (such as  $\text{CO}_2$  dissolved in soda to produce carbonic acid) is a two-phase system
- Two liquids, such as ethanol and water is a single phase system

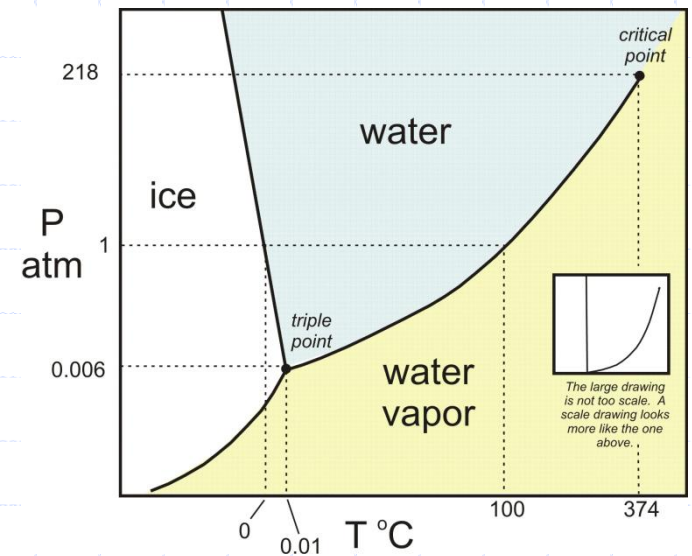


## 17.1.3 – Phase Diagrams

- A **phase diagram** is a graphical plot of pressure vs temperature.
- Has three regions: solid, liquid, gas
- **Triple point** is where the three boundary lines intersect



A typical phase diagram for a pure substance



The phase diagram for water ( $\text{H}_2\text{O}$ ) is different, why?

